3. STRUCTURE OF ATOM

Up to the end of 19^{th} century, the Dalton's atomic model remained undisputed. But the new discoveries towards the end of 19^{th} century and early 20^{th} century showed that atom has a complex structure and is not indivisible. The studies also showed that the atoms are made up electrons, protons and neutrons. Properties of elements are explained on the basis of negatively charged electrons around a positively charged mass, situated at the centre of the atom, called nucleus. The positively charged nucleus containing protons and neutrons occupies much less space in an atom compared to the large space in which the electrons are distributed.

Table

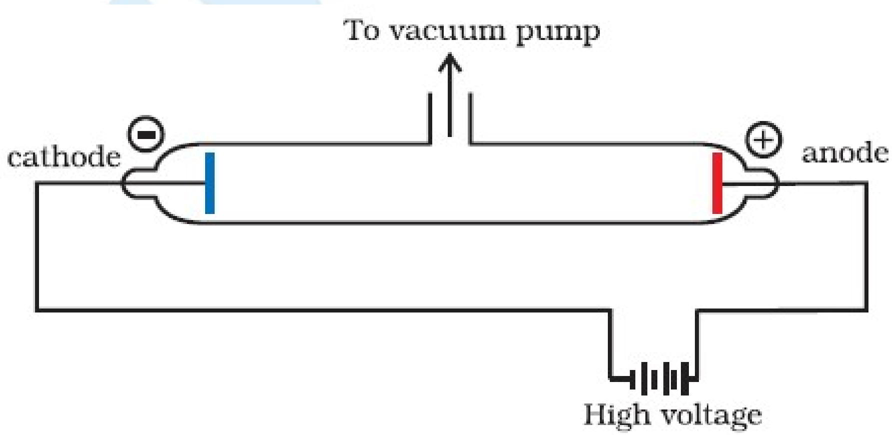
Positions, relative masses and relative charges of electrons, protons and neutrons

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg	Mass/u	Approx. mass/u
Electron	e	-1.6022×10^{-19}	-1	9.10939×10^{-31}	0.00054	0
Proton	p	$+1.6022\times10^{-19}$	+1	1.67262×10^{-27}	1.00727	1
Neutron	n	0	0	1.67493×10^{-27}	1.00867	1

ELECTRONS

After the Dalton's atomic model, one of the significant advances was made by J.J. Thomson. This was the discovery of charged particles called electrons. They are extremely small in size. The existence of electrons is proved by experiments using discharge tubes. When an electric discharge at a very high voltage is passed through a gas contained in a glass tube at a very low pressure, a particular type of rays emanate from the cathode which are known as cathode rays. These rays travel in straight lines. The particles constituting these rays possess significant amounts of kinetic energy. When these rays are allowed to pass through a strong electric field, they are deflected away from the negative plate. These facts lead to the conclusion that cathode rays consist of rapidly moving negatively charged particles. These particles are known as electrons. They are present in all atoms around the nucleus in definite energy levels.

The absolute mass of electron is $9.1\text{x}10^{-31}$ kg which is approximately 1/1840 of the mass of a proton or a hydrogen atom. The absolute charge of electron is 1.6t02 x 10^{-19} coulomb of negative charge. This has been taken to be one unit of negative charge.



PROTONS

Like electrons, protons also present in all atoms. Presence of protons was observed by Goldstein in 1896 in a discharge tube with a perforated cathode. It was observed that a stream of protons passed through the holes in the cathode in a direction away from the anode. A beam of protons (or any other positively charged particles) is also known as positive rays.

Experiments carried out by Lord Rutherford showed that positive particles produced in a discharge tube containing hydrogen gas are same as protons.

Protons are positively charged particles. Mass of each proton is same as that of a hydrogen atom. But the mass of hydrogen atom is 1 a.m.u., therefore the relative mass of a proton is 1 a.m.u. However, the absolute mass of a proton is 1.6×10^{-27} kg. The charge of proton is equal and opposite to the charge of an electron. So, the absolute charge of a proton is 1.6×10^{-19} coulomb of positive charge which has been taken to be one unit of positive charge.

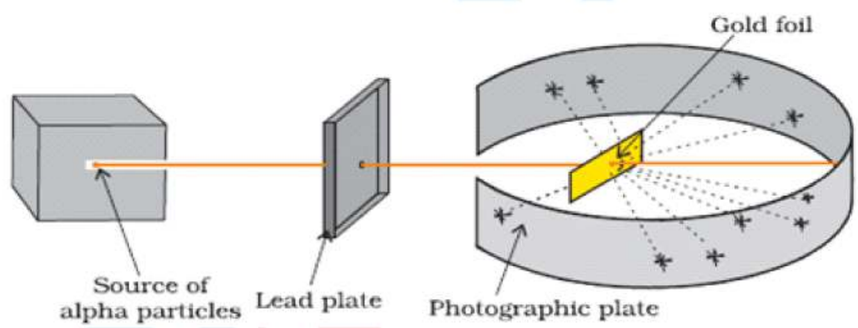
NEUTRONS

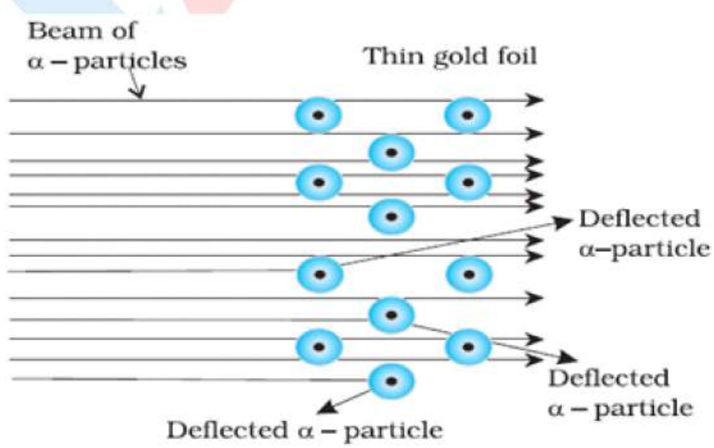
One atom of hydrogen has one proton and one electron. One atom of helium has two protons and two electrons. Therefore, the relative atomic mass of helium should be twice that of hydrogen. But in fact the relative atomic mass of helium in 4 and not 2.

This discrepancy was removed by the discovery of neutrons by James Chadwick in 1932. Chadwick's experiments showed that apart from protons, nuclei of all atoms except that of hydrogen atom, contain another type of particles which have the same mass as hydrogen atoms or protons. These particles do not carry any electric charge i.e., they are electrically neutral. Thus an atom of helium contains 2 protons and 2 neutrons which make its relative mass equal to 4.

RUTHERFORD'S EXPERIMENT–ATOMIC MODEL

In 1911, Lord Rutherford bombarded a gold leaf with a stream of alpha-particles (an alpha-particle is a positively neutrons). He observed that whereas most of the alpha-particles penetrated the gold leaf, some of them were deflected from their path; some alpha-particles were even repelled by the gold leaf.





Concluding the results of his experiment, Rutherford suggested that:

1. The central part of an atom, called nucleus, is positively charged;
2. Nearly total mass of the atom is concentrated in the nucleus of the atom;
3. The space occupied by the central part (nucleus) of the atom is much less than the much larger space in which the electrons revolve around the nucleus.

Based on the above experiment, Rutherford proposed that each atom consisted of a positive nucleus around which negatively charged electrons are revolving. Since an atom is electrically neutral, therefore, the number of electrons in the atom of an element should be the same as the number of protons in its nucleus. Accordingly, in an atom of hydrogen, the lightest element, helium, two electrons and two protons are present.

ATOMIC NUMBER AND MASS NUMBER

Atomic Number. The number of positive charge carried by the nucleus of an atom is termed as atomic number or charge number (Moseley 1913) is represented by symbol Z. Since the charge on the nucleus is equal to the number of the protons (P) in it which in turn is equal to the number of electrons (e) in the atom, the atomic number is defined as the number of protons or the number of electrons for a neutral atom in the atom.

Mass Number. The total number of the protons and neutrons present in the nucleus of the atom is known as the mass number of the atom.

BOHR'S MODEL OF AN ATOM

In 1913, Niels Bohr put forward his theory to explain the structure of an atom. According to Bohr's theory:

1. The nucleus of an atom is situated at its centre.
2. The electrons in an atom revolve around the nucleus in definite circular paths known as energy levels.
3. Each energy level has a fixed amount of energy.
4. The energy levels are either designated at K,L,M, N, etc. or numbered n = 1, 2, 3, 4 etc. outwards from the nucleus.
5. The change in the energy of an electron takes place only when it jumps from a higher energy level to a lower energy level (loss of energy), or, when it jumps from a lower energy level to a higher energy level (gain of energy).
6. Thus, as long as electrons continue to revolve in the same energy level, they neither lose nor gain energy and the atom remains stable.

DISTRIBUTION OF ELECTRONS IN DIFFERENT ENERGY LEVELS

The distribution of electrons in different energy level is governed by a scheme known as Bohr-Bury scheme which states that:

(i) The maximum number of electrons that can be accommodated in any energy level in 2n2 where n is number of that energy level. The energy levels are also known as shells. Thus,

(ii) K- shell (n=1) can have $2x1^2 = 2$ electrons.

(iii) L- shell (n=2) can have $2x\ 2^2 = 18$ electrons.

(iv) M- shell (n=3) can have $2x3^2 = 18$ electrons.

(v) N- shell (n=4) can have $2x4^2 = 32$ electrons.

(vi) The electrons first occupy the shell with the lower energy.

(vii) The outermost shell of an atom cannot have more than 8 electrons and next to the outermost shell cannot have more than 18 electrons.

The systematic distribution of electrons in different energy shell is called the electronic configuration of the atom.

VALENCE SHELL AND VALENCE ELECTRONS

The outermost orbit of an atom is called its valence shell. The electrons present in the outermost shell of an atom are known as the valence electrons. Only valence electrons of an atom take part in chemical changes and determine its combining capacity which is known as valency of the atom.

ISOTOPES AND ISOBARS

Istopes. In 1919, F.W. Aston discovered that the atoms of some naturally-occuring elements were not exactly alike. He observed that these atoms of the same element had different masses. Such types of atoms of the same element are known as isotopes. All the isotopes of any element have the same atomic number because their nuclei contain the same number of protons, but their mass numbers are different because the numbers of neutrons in their nuclei are different. Since the isotopes of any element contain the same number of electrons, therefore, they have the same chemical properties. However, as the number of neutrons in the nuclei of the atoms of different isotopes is different, their physical properties like densities and melting and boiling points are slightly different. Examples of isotopes of some elements are:

Hydrogen – Protium ($^{1}_{1}H$), Deuterium, Tritium ($^{3}_{1}H$)

Chlorine – Chlorine-35 ($^{35}_{17}Cl$), Chlorine-37 ($^{37}_{17}Cl$)

Carbon – Carbon-12 ($^{12}_{6}C$), Carbon -14 ($^{14}_{6}C$)

Isobars. It is a well known fact that the atomic number of no two elements is same. In general, the mass numbers of the different elements are also different. However, there are some cases, like argon ($^{40}_{18}Ar$) and calcium ($^{40}_{20}Ca$), where the mass number is the same, and as the elements are different, the atomic numbers are different. Such pairs of elements are known as isobars. Hence, isobars may be defined as those elements which have the same mass number (atomic mass) but different atomic numbers.